REVIEW QUESTIONS
Chapter 8

Use only a periodic table to answer the following questions.

1. Write complete electron configuration for each of the following elements:

   a) Aluminum (Al)

      $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

   b) Sulfur (S)

      $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

   c) Manganese (Mn)

      $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^5$

2. Write condensed orbital diagrams for each of the following elements and determine the number of unpaired electrons in each:

   a) Zinc (Zn)

      [Ar] ↑↓ (4s) ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ (3d) (0 unpaired)

   b) Selenium (Se)

      [Ar] ↑↓ (4s) ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ (3d) ↑↓ ↑ ↑ (4p) (2 unpaired)

   c) Lead (Pb) **(5d and 4f orbital are omitted for clarity)**

      [Xe] ↑↓ (6s) ↑ ↑ __ (6p) (2 unpaired)

3. Identify the element that belongs to each of the following electron configurations:

   a) $1s^2\ 2s^2\ 2p^6\ 3s^2$ **Mg**

   b) $[Ne]\ 3s^2\ 3p^1$ **Al**

   c) $[Ar]\ 4s^1\ 3d^5$ **Cr**

   d) $[Kr]\ 5s^2\ 4d^{10}\ 5p^4$ **Te**

4. A section of the periodic table with all identification features removed is shown below.

V	W	X
	Y	Z

Which element has the smallest atomic radius? Give a brief explanation for your choice.

X would be expected to have the smallest radius. This is so because moving left to right across a period increases the effective nuclear charge (Z_{eff}) due to addition of protons, and therefore decreases atomic radius. Furthermore, moving down a group increases atomic radius due to additional energy levels. Therefore the smallest atom would be in the upper right hand corner of periodic table.

5. A period 3 element has the following ionization energies. What is the identity of this element?

IE_1= 801 kJ/mol
IE_2= 2427 kJ/mol
IE_3= 3660 kJ/mol
IE_4= 25025 kJ/mol

Since there is a very large increase in IE_4 compared to IE_3, this element would be expected to have 3 valence electrons. Aluminum is the element on the third period with 3 valence electrons.

6. Arrange each of the following elements in order of increasing atomic radius.

a) F, P, S, As

F < S < P < As

b) B, Ca, Ga, Cs

B < Ga < Ca < Cs

c) Na, Al, P, Cl, Mg

Cl < P < Al < Mg < Na

7. Arrange the following in order of increasing first ionization energy.

a) Na, Cl, Al, S, Cs

Cs < Na < Al < S < Cl

b) F, K, P, Ca, Ne

K < Ca < P < F < Ne

c) Ne, Na, P, Ar, K

K < Na < P < Ar < Ne

8. The first four ionization energies of yttrium (Z=39) are IE_1= 616, IE_2= 1180, IE_3= 1980, IE_4= 5960 kJ/mol. Answer the following questions based on these data:

a) Explain the increasing trend in the successive energies of yttrium.

Successive ionization energies are always greater because the electron is successively removed from a more positive ion each time.

b) Explain the large increase in IE_2 compared to IE_1.

The second electron in yttrium removed during IE_2 is a completely filled 5s subshell and more stable than the first electron removed during IE_1.

↑↓ ↑ __ __ __ __.
5s 4d

c) Explain the large increase in IE_4 compared to IE_3.

The fourth electron removed during IE_4 is a core electron from a noble gas configuration and therefore very stable, making IE_4 very large.

9. Explain why alkali metals have a greater electron affinity than alkaline earth-metals.

Alkaline earth-metals have a complete filled s subshell and therefore do not desire electrons very readily. Alkali metals on the other hand need one electron to complete their s subshell and would readily accept an electron, therefore leading to large electron affinities.

10. Which element in each of the following sets would you expect to have the highest second ionization energy (IE_2)?

a) Na, K, Fe

For the second ionization energy, both Na and K remove electrons from the core electrons with the noble gas configuration, whereas Fe would remove the 4s electron. Na is smaller than K, therefore it would have the largest IE_2 of the three elements.

b) Na, Mg, Al

The second ionization energy of Na removes an electron from the core shell with a noble gas configuration while Mg and Al lose 3s electrons during IE_2. Therefore Na would have the largest IE_2 value.

11. Until the early 1960s the group 8A elements were called inert gases. They are no longer referred to as such, since Xe and Kr were found to react with some substances. Suggest a reason why Xe would react with fluorine, but Ne would not.

Xe is a much larger atom than Ne, therefore it would have a much lower IE than Ne. As a result, Xe could react with elements with large electron affinities such as F.

12. The table below gives the electron affinities in kJ/mol for group 1B and 2B elements.

1B	2B
Cu –119	Zn >0
Ag –126	Cd >0
Au –223	Hg >0

a) Explain why group 1B elements have negative electron affinities, while group 2B elements have positive values.

Group 2B elements have a complete d orbital and therefore would not be very eager to add additional electrons to their shells. Therefore they would be expected to have positive EA. Group 1B elements, however, have an incomplete s subshell which would readily accept an additional electron. Therefore they would have negative EA.

b) Explain why group 1B electron affinities become more negative moving down the group.

Group 1B elements have the generic ns^1 $(n–1)d^{10}$ electron configuration and when accepting additional electrons would experience inter-electron repulsions between the electrons on the same energy level. Going down the group, the size of the atom increases and therefore reduces this inter-electron repulsion. As a result, the addition of the electron is more favored leading to a larger EA.

13. For each set shown below, select the atoms or ions that are isoelectronic with each other, and write their electron configuration:

a) K^+, Rb^+, Ca^{2+}

K^+ and Ca^{2+} are isoelectronic [Ne] $3s^2$ $3p^6$

b) S^{2-}, Ar, Se^{2-}

S^{2-} and Ar are isoelectronic [Ne] $3s^2$ $3p^6$

c) Mg^{2+}, Cl^-, Al^{3+}

Mg^{2+} and Al^{3+} are isoelectronic $1s^2$ $2s^2$ $2p^6$

14. Explain each of the following trends in ionic radii:

a) $I^- > I > I^+$

All species have 5 energy levels. I^- has the largest number of electrons (54) and has the largest inter-electron repulsion, making it the largest. I^+ has the lowest number of electrons (52) and has the lowest inter-electron repulsion, making it the smallest.

b) $Ca^{2+} > Mg^{2+} > Be^{2+}$

Ca^{2+} has 3 energy levels, Mg^{2+} has 2 energy levels, while Be^{2+} has only one energy level. Therefore Ca^{2+} is the largest and Be^{2+} is the smallest.

15. Arrange the atoms or ions in each of the following sets in order of increasing radius:

a) Br^-, Na^+, Mg^{2+}

$Mg^{2+} < Na^+ < Br^-$

b) Ar, Cl^-, S^{2-}

$Ar < Cl^- < S^{2-}$

c) Co^{3+}, Fe^{2+}, Fe^{3+}

$Co^{3+} < Fe^{3+} < Fe^{2+}$

d) K^+, Cl^-, Ca^{2+}, P^{3-}

$Ca^{2+} < K^+ < Cl^- < P^{3-}$

In an isoelectronic series size increases as the Z_{eff} decrease (charge becomes less positive)